PHAR 123 GENERAL CHEMISTRY for HEALTH SCIENCES

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General Chemistry for Health Sciences PHAR 123

2020-2021 FALL TERM

Asst. Prof. Behiye ÖZTÜRK ŞEN

İSTANBUL ALTINBAŞ ÜNİVERSİTESİ

General Chemistry for Health Sciences
PHAR 123

Wednesday:9:45-11:30
Thursday:9:45-11:30

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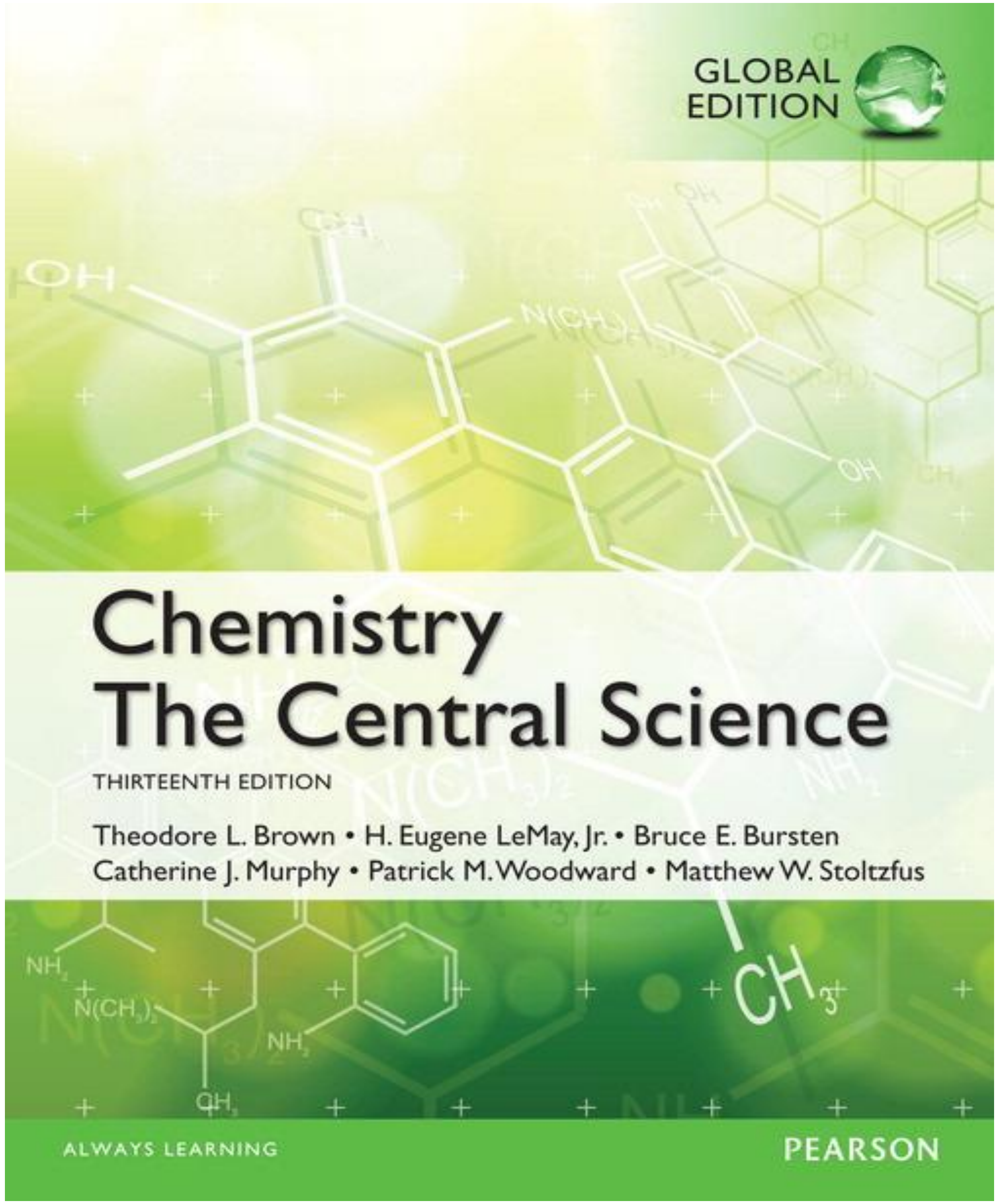
COURSE CONTENT		
Week	Topics	Study Materials
1	Matter and Measurements	Text Book
2	Atoms, Molecules, and Ions	Text Book
3	Stoichiometry	Text Book
4	Aqueous Solutions and Solution Chemistry	Text Book
5	Thermochemistry	Text Book
6	MIDTERM I	
7	Electronic Structure of Atoms	Text Book
8	Periodic Properties of the Elements	Text Book
9	Chemical Bonding	Text Book
10	Gases	Text Book
11	MIDTERM II	
12	Intermolecular Forces, Liquids and Solids	Text Book
13	Properties of Solutions	Text Book
14	Chemical Kinetics and Chemical Equilibrium	Text Book
15	FINAL	

RECOMMENDED SOURCES	
Textbook	Chemistry: The Central Science, 13th ed., Theodore L. Brown, H. Eugene LeMay, Jr.Bruce E. Bursten, Chatherine J. Murphy, Patrick Woodward, Pearson/Prentice Hall, 2014
Additional Resources	General Chemistry: Principles and Modern Application, 10th ed., Ralph H. Petrucci, William S. Harwood, F. Geoffrey Herring, Prentice Hall, Pearson Educational International, 2010

ASSESSMENT		
IN-TERM STUDIES	QUANTITY	PERCENTAGE
Homeworkl	1	40
Quiz	2	30
Final	1	30
Total	3	100



Lecture Presentation

Chapter 1

Introduction: Matter and Measurement

James F. Kirby
Quinnipiac University
Hamden, CT

In this Chapter;
You will learn

1. **Chemistry**
2. **Matter**
3. Classification of Matter According to States (Solid,Liquid and Gas) and Composition (Element, Compound, Mixture)
4. **Types of Properties**:

- ✓ Physical Properties (Boiling Points, density)
- ✓ Chemical Properties (Flammability, Corrosiveness)
- ✓ Intensive Property
- ✓ Extensive Property

5. **Types of Changes:**

- ✓ Physical Changes (Temperature and Volume)
- ✓ Chemical Changes (Combustion, Oxidation and Decomposition)

In this Chapter;
You will learn

6. **Numbers and Chemistry (Units of Measurement)**
7. **Uncertainities in Measurement (Exact Number and Inexact Number)**
8. **Significant Figures (Uncertainity)**
9. **Significant Figures in Calculations**

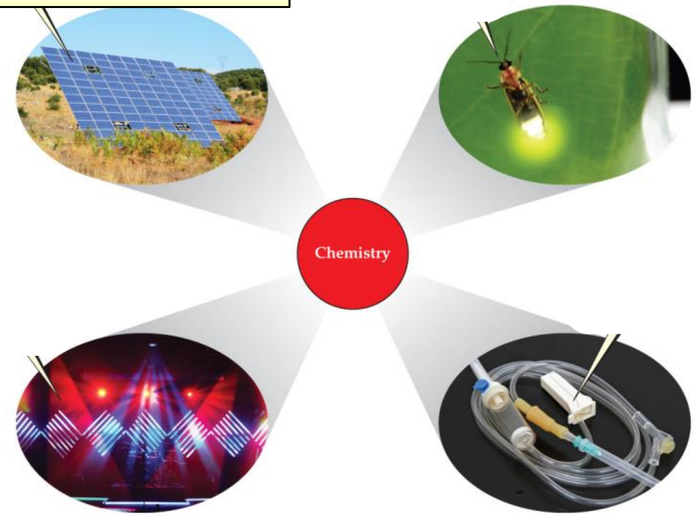
- ✓ Addition and Substraction
- ✓ Multiplication and Division

10. **Rounding of Numbers**
11. **Dimensional Analysis**

1.Chemistry

Energy: Solar panels are composed of specially treated silicon

Biochemistry
The flash of the firefly results from a chemical reaction in the insect

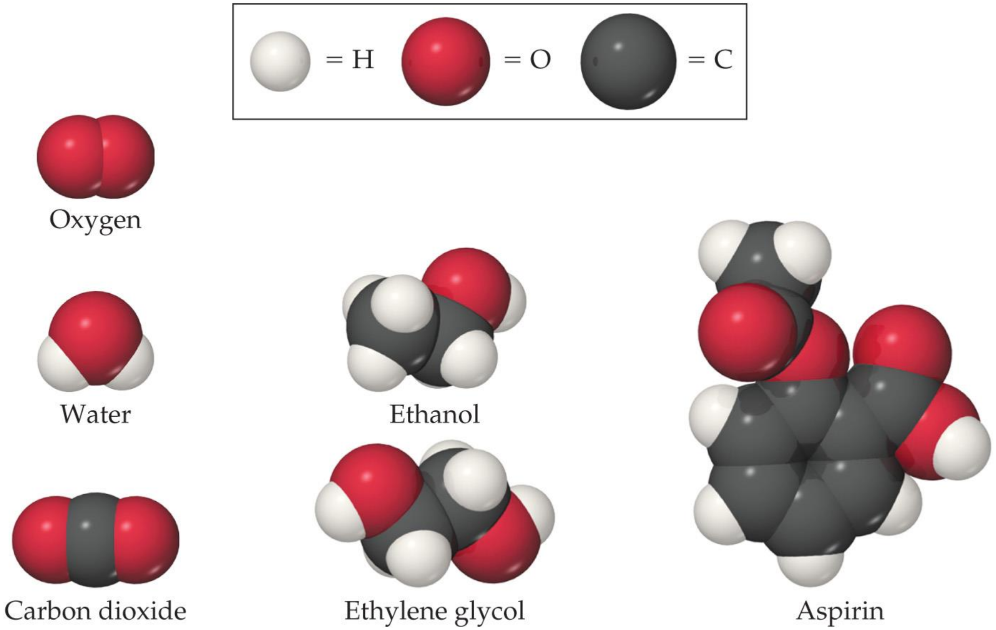


Technology: LEDs light emitting diodes are formed from elements such as gallium, arsenic and phosphorus

Medicine: Connectors and tubing for medical procedures such as intravenous injections are made from plastics highly resistant to chemical attack

- Chemistry is the study of the **properties** and **behavior** of **matter**.
- It is central to our fundamental understanding of many science-related fields.
- All matter is comprised of combinations of only ~ **100 substances** called elements
- **Goal:** relate to the properties of the matter to the particular element it contains.

2. Matter



- **Atoms** are the building blocks of matter.
- Each **element** is made of a unique kind of atom.
- A **compound** is made of two or more different kinds of elements.

Observable changes in matter: macroscopically

Behaviour of atoms and molecules: Submicroscopic level

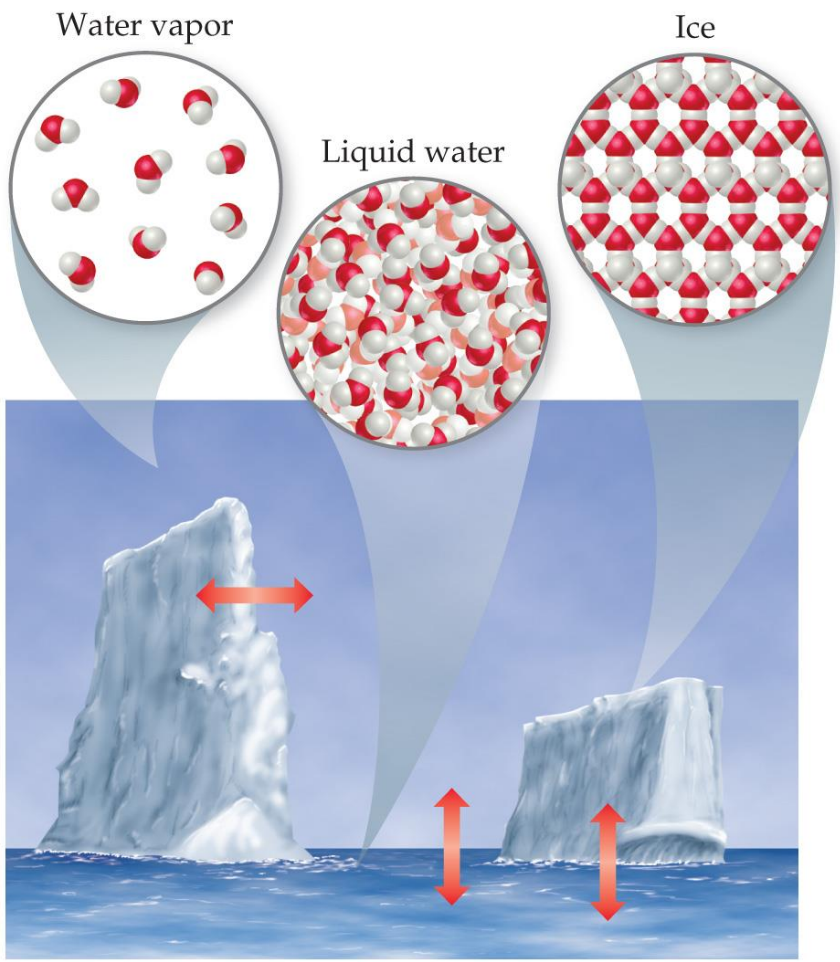
Chemistry : the science that seeks to understand behavior of atoms and molecules.

3.Classification of Matter

States
- Gas
- Liqiud
- Solid

Composition
- Element
- Compound
- Mixture

States of Matter



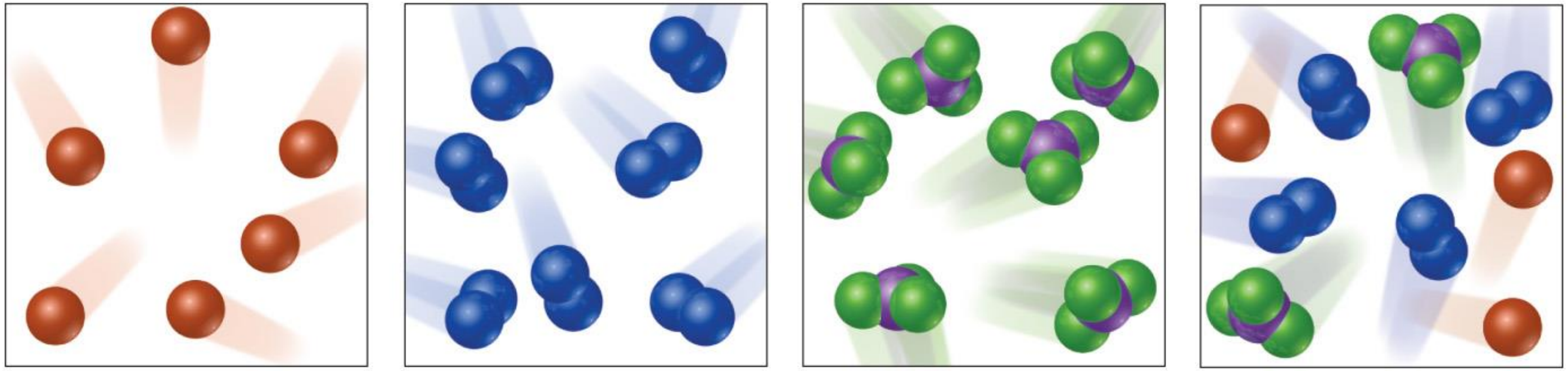
- The three states of matter are
  1) solid.
  2) liquid.
  3) gas.
- In this figure, those states are ice, liquid water, and water vapor.
- Volume & shape
- Movements of molecules
- Inter-molecular distances

Classification of Matter - Substances

- A **substance** has distinct properties and a composition that does not vary from sample to sample. **Ex:** Table salt and water
- The two types of substances are **elements** and **compounds**.

1. An **element** is a substance which **can *not*** be decomposed to simpler substances.
2. A **compound** is a substance which ***can* be** decomposed to simpler substances.

Classification of Matter - Substances



a) Atoms of an element

b) Molecules of an element

c) Molecules of a compound

d) Mixtures of elements and a compound

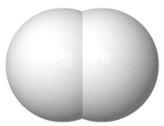
Only one kind of atom is in any element

Compounds must have at least two kinds of atoms

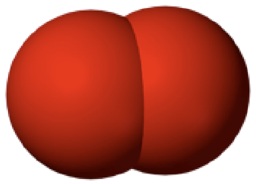
Mixtures: atoms and molecules retain their chemical identity
118 elements: earths crust' atmosphere, human body
Periodic table: symbols and names of elements

Compounds and Composition

- Compounds have a definite composition. That means that the relative number of atoms of each element that makes up the compound is the same in any sample. ***Regardless of source: nature or lab!***
- This is **The Law of Constant Composition** (or **The Law of Definite Proportions**). ***Unique characteristics!***



Hydrogen molecule (H_2)



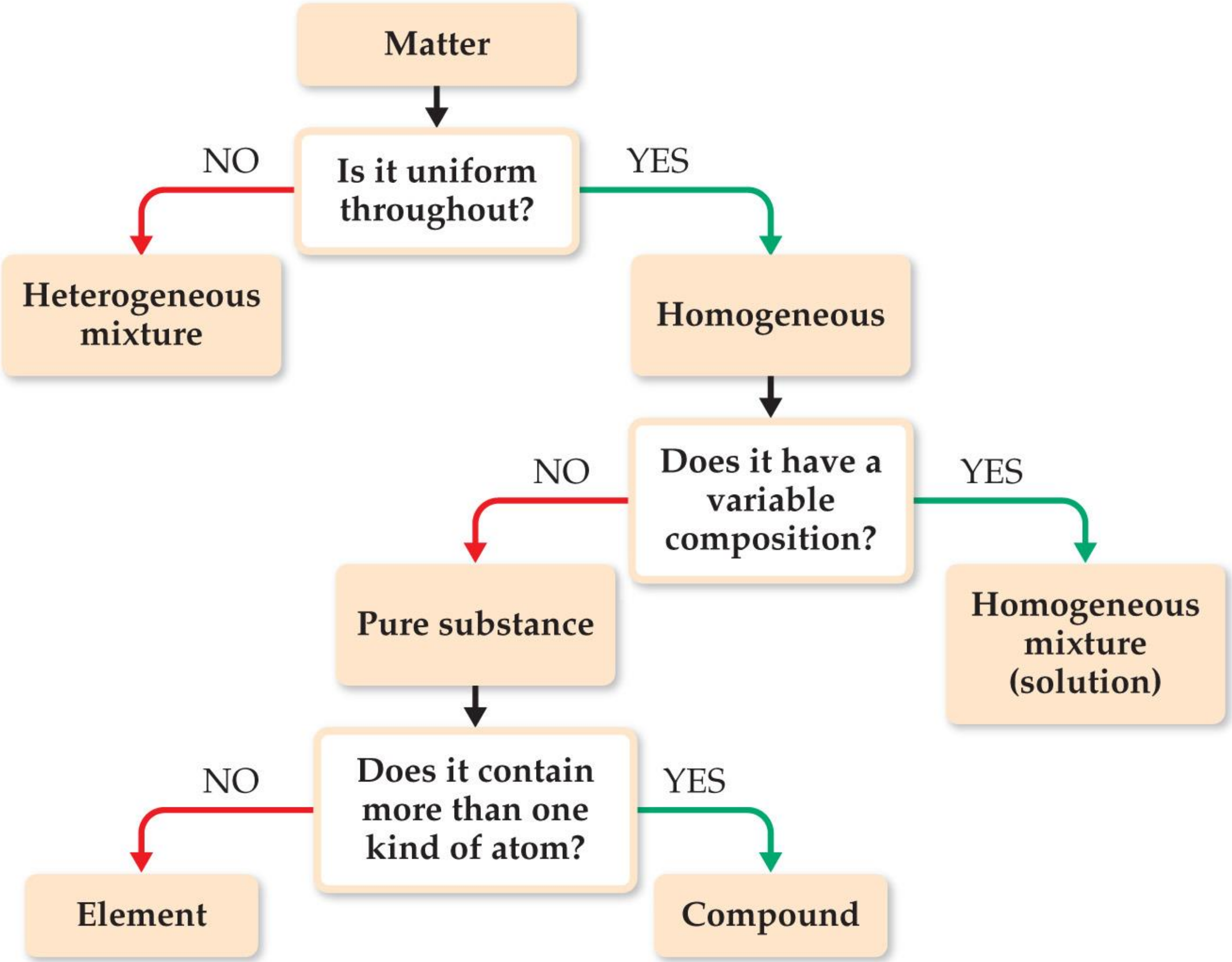
Oxygen molecule (O_2)

The elements hydrogen and oxygen themselves exist naturally as diatomic molecules

Classification of Matter - Mixtures

- **Matter in real world:** mixtures
- **Mixtures** exhibit the properties of the substances that make them up
- Mixtures do not have fixed ratio: coffee with sugar
- Mixtures can vary in composition throughout a sample **(heterogeneous)** or can have the same composition throughout the sample **(homogeneous)**
- Another name for a homogeneous mixture is **solution**
- **Component:** substances making up mixtures. Elements and molecules retain their chemical identity in a mixture
- Elements and molecules retain their chemical identity in a mixture

Classification of Matter Based on Composition



- ❖ If you follow this scheme, you can determine how to classify any type of matter.
- ➢ Homogeneous mixture
- ➢ Heterogeneous mixture
- ➢ Element
- ➢ Compound

Separating Mixtures

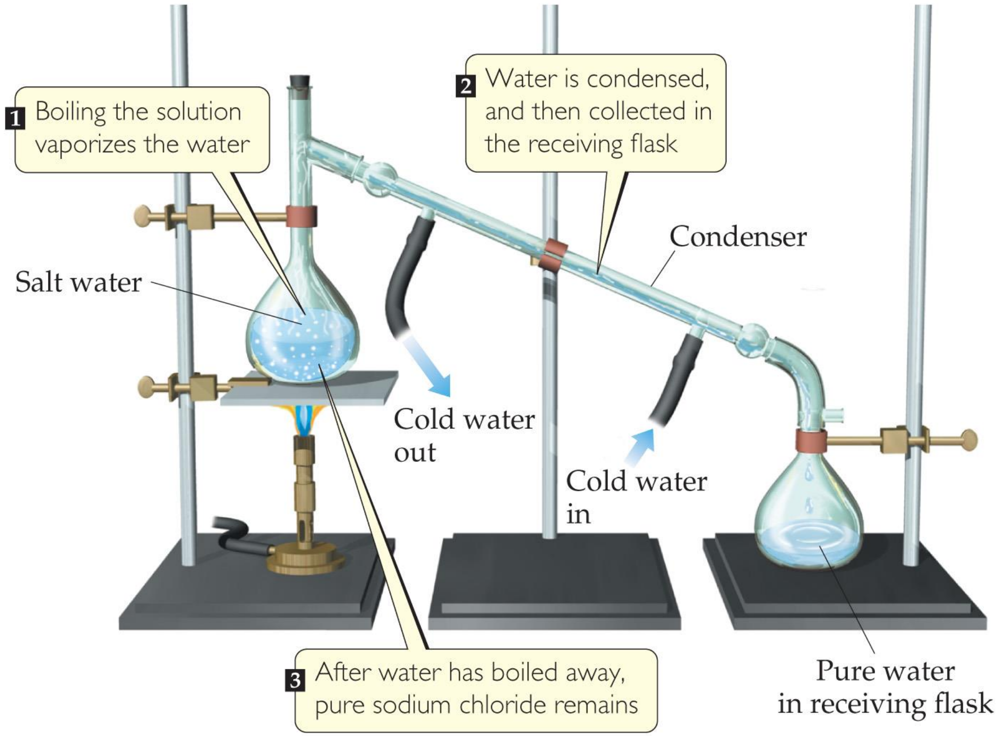
- Sands in a bucket of water or sands and iron
- Mixtures can be separated based on physical properties of the components of the mixture. Some methods used are

  - ➢ filtration
  - ➢ distillation
  - ➢ chromatography

Filtration

- In filtration, solid substances are separated from liquids and solutions.

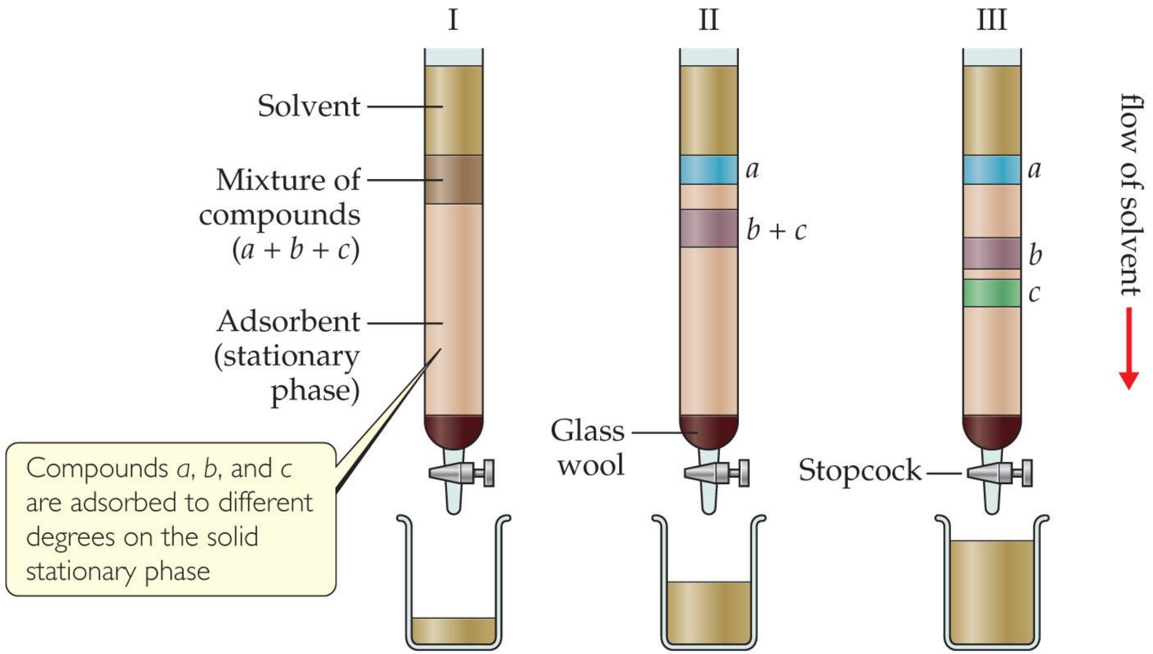
Distillation



- Distillation uses differences in the boiling points of substances to separate a homogeneous mixture into its components.

Chromatography

- This technique separates substances on the basis of differences in the ability of substances to adhere to the solid surface, in this case, dyes to paper.



4. Types of Properties

- Substances have unique properties
- **Physical Properties** can be observed without changing a substance into another substance.
  - Some examples include ***boiling point***, ***density***, ***mass***, or ***volume***.
- **Chemical Properties** can *only* be observed when a substance is changed into another substance.
  - Some examples include flammability, ***corrosiveness***, or ***reactivity*** with acid.

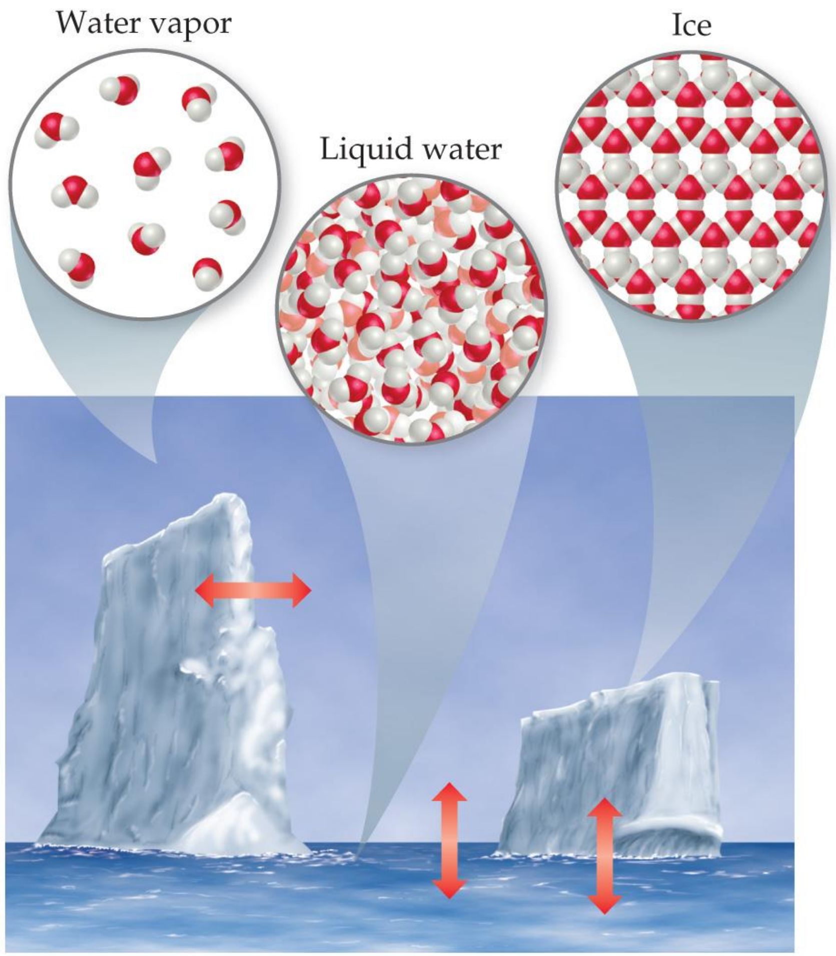
4. Types of Properties

- **Intensive Properties** are independent of the amount of the substance that is present.
  - *Examples include density, boiling point, or color.*
- **Extensive Properties** depend upon the amount of the substance present.
  - *Examples include mass, volume, or energy.*

5. Types of Changes

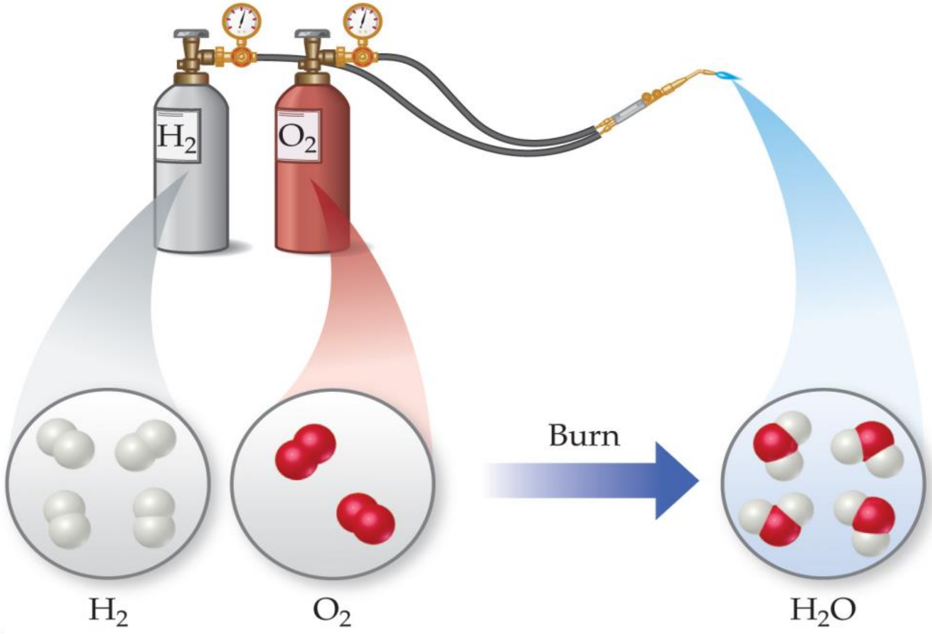
- **Physical Changes** are changes in matter that do *not* change the composition of a substance.
  - Examples include changes of state, ***temperature***, and ***volume***.
- **Chemical Changes** result in new substances.
  - Examples include ***combustion***, ***oxidation***, and ***decomposition***.

Changes in State of Matter



- Converting between the three states of matter is a **physical change**.
- When ice melts or water evaporates, there are still 2 H atoms and 1 O atom in each molecule.

Chemical Reactions (Chemical Change)



In the course of a chemical reaction, the reacting substances are converted to new substances. Here, the elements hydrogen and oxygen become water.

Give it Some Thought

- Which of these changes are physical and which are chemical?

a) Plants make sugar from carbondioxide and water
b) Water vapor in the air forms frost
c) A goldsmith melts a nugget of gold and pulls it into a wire

6. Numbers and Chemistry

- Numbers play a major role in chemistry. Many topics are **quantitative** (have a numerical value).
- Concepts of numbers in science
  - Units of measurement
  - Quantities that are measured and calculated
  - Uncertainty in measurement
  - Significant figures
  - Dimensional analysis

Units of Measurement—Metric System

- If the number represents a measured quantity, unit should also be specified along with the number.
- Metric system is used for units of scientific measurements: France vs English System
- The base units used in the metric system
  - Mass: gram (g)
  - Length: meter (m)
  - Time: second (s or sec)
  - Temperature: degrees Celsius (^{o}C) *or* Kelvins (K)
  - Amount of a substance: mole (mol)
  - Volume: cubic centimeter (cc or cm^3) *or* liter (l)

Units of Measurements—SI Units

SI units : particular choice of metric system

Table 1.4 SI Base Units

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s or sec
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A or amp
Luminous intensity	Candela	cd

- *Système International d'Unités* ("The International System of Units")
- A different base unit is used for each quantity.

Units of Measurement - Metric System Prefixes

Table 1.5 Prefixes Used in the Metric System and with SI Units

Prefix	Abbreviation	Meaning	Example	
Peta	P	10^{15}	1 petawatt (PW)	$= 1 \times 10^{15}$ watts[a]
Tera	T	10^{12}	1 terawatt (TW)	$= 1 \times 10^{12}$ watts
Giga	G	10^{9}	1 gigawatt (GW)	$= 1 \times 10^{9}$ watts
Mega	M	10^{6}	1 megawatt (MW)	$= 1 \times 10^{6}$ watts
Kilo	k	10^{3}	1 kilowatt (kW)	$= 1 \times 10^{3}$ watts
Deci	d	10^{-1}	1 deciwatt (dW)	$= 1 \times 10^{-1}$ watt
Centi	c	10^{-2}	1 centiwatt (cW)	$= 1 \times 10^{-2}$ watt
Milli	m	10^{-3}	1 milliwatt (mW)	$= 1 \times 10^{-3}$ watt
Micro	μ[b]	10^{-6}	1 microwatt (μW)	$= 1 \times 10^{-6}$ watt
Nano	n	10^{-9}	1 nanowatt (nW)	$= 1 \times 10^{-9}$ watt
Pico	p	10^{-12}	1 picowatt (pW)	$= 1 \times 10^{-12}$ watt
Femto	f	10^{-15}	1 femtowatt (fW)	$= 1 \times 10^{-15}$ watt
Atto	a	10^{-18}	1 attowatt (aW)	$= 1 \times 10^{-18}$ watt
Zepto	z	10^{-21}	1 zeptowatt (zW)	$= 1 \times 10^{-21}$ watt

[a]The watt (W) is the SI unit of power, which is the rate at which energy is either generated or consumed. The SI unit of energy is the joule (J); $1\ J = 1\ kg \cdot m^2/s^2$ and $1\ W = 1\ J/s$.
[b]Greek letter mu, pronounced "mew."

- ***Prefixes*** convert the base units into units that are appropriate for common usage or appropriate measure.
- Milli: 10^{-3} fraction of a unit
- Problem solving
- Exponential

Mass and Length

- These are basic units we measure in science.
- **Length** is a measure of distance. The meter is the base unit.
- **Mass** is a measure of the amount of material in an object. SI uses the kilogram as the base unit. The metric system uses the gram as the base unit. Other units of mass is obtained by adding prefixes to the word of gram.

Temperature

- In general usage, **temperature** is considered the “hotness and coldness” of an object that determines the direction of heat flow.
- Heat flows spontaneously from an object with a higher temperature to an object with a lower temperature.
- In scientific measurements, the **Celsius** and **Kelvin** scales are most often used.

Temperature

The **Celsius scale** is based on the properties of water.

- 0 °C is the freezing point of water.
- 100 °C is the boiling point of water

The **Kelvin** is the SI unit of temperature.

- It is based on the properties of gases.
- There are no negative Kelvin temperatures.
- The lowest possible temperature is called absolute zero (0 K).

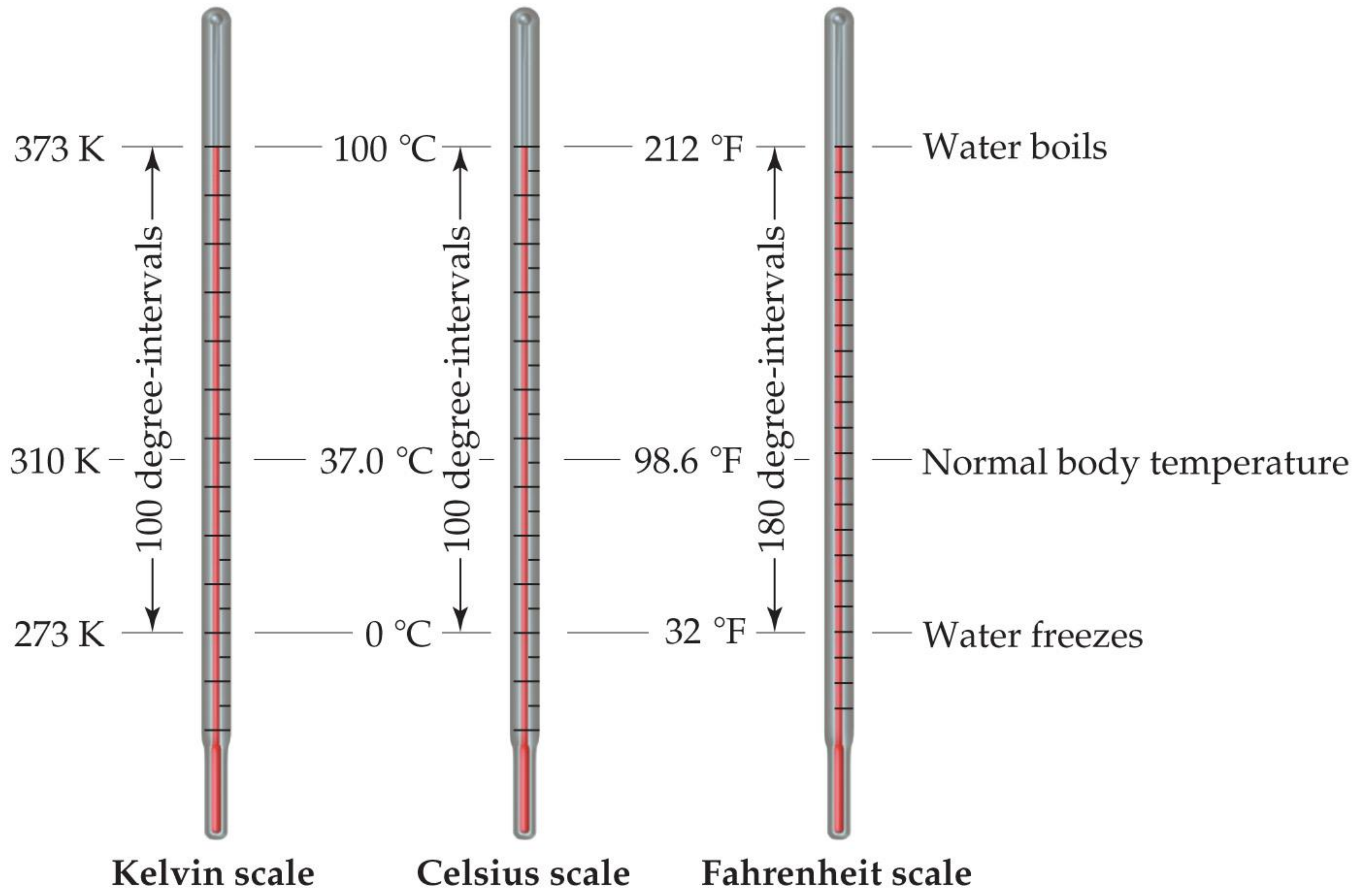
K = °C + 273.15

Temperature

The Fahrenheit scale is not used in scientific measurements, but you hear about it in weather reports!

The equations below allow for conversion between the Fahrenheit and Celsius scales:

- **$°F = 9/5(°C) + 32$**
- **$°C = 5/9(°F - 32)$**

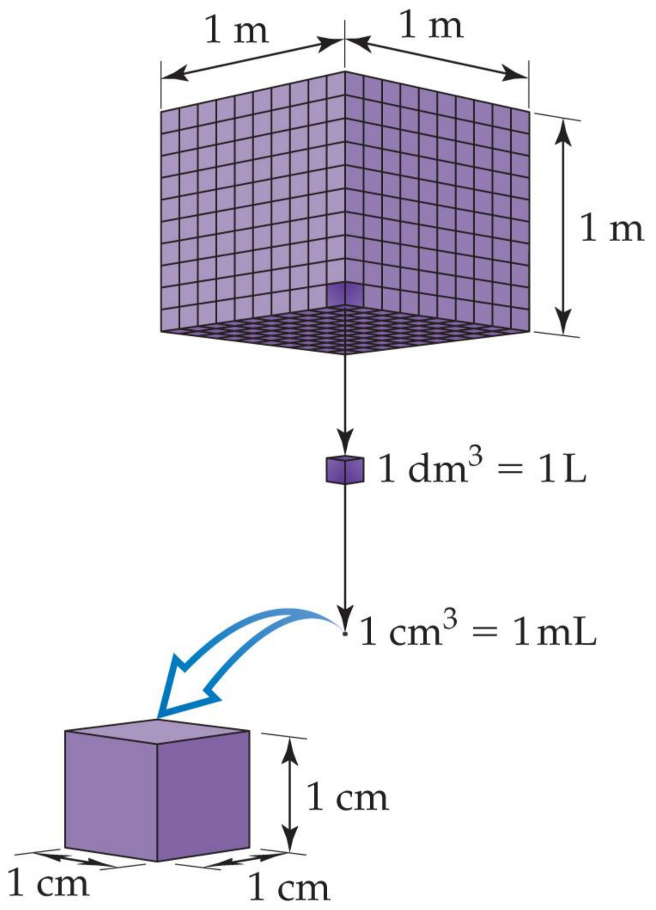

373 K
310 K
273 K
100 degree-intervals
100 °C
37.0 °C
0 °C
100 degree-intervals
212 °F
98.6 °F
32 °F
180 degree-intervals
Water boils
Normal body temperature
Water freezes
Kelvin scale
Celsius scale
Fahrenheit scale

Derived SI units - Volume

- Note that volume is not a base unit for SI; it is derived from length (m × m × m = m^3). Multiplication
- The most commonly used metric units for volume are the liter (L) and the milliliter (mL).

✓ A liter is a cube 1 decimeter (dm) long on each side.

✓ A milliliter is a cube 1 centimeter (cm) long on each side, also called 1 cubic centimeter (cm × cm × cm = cm^3).

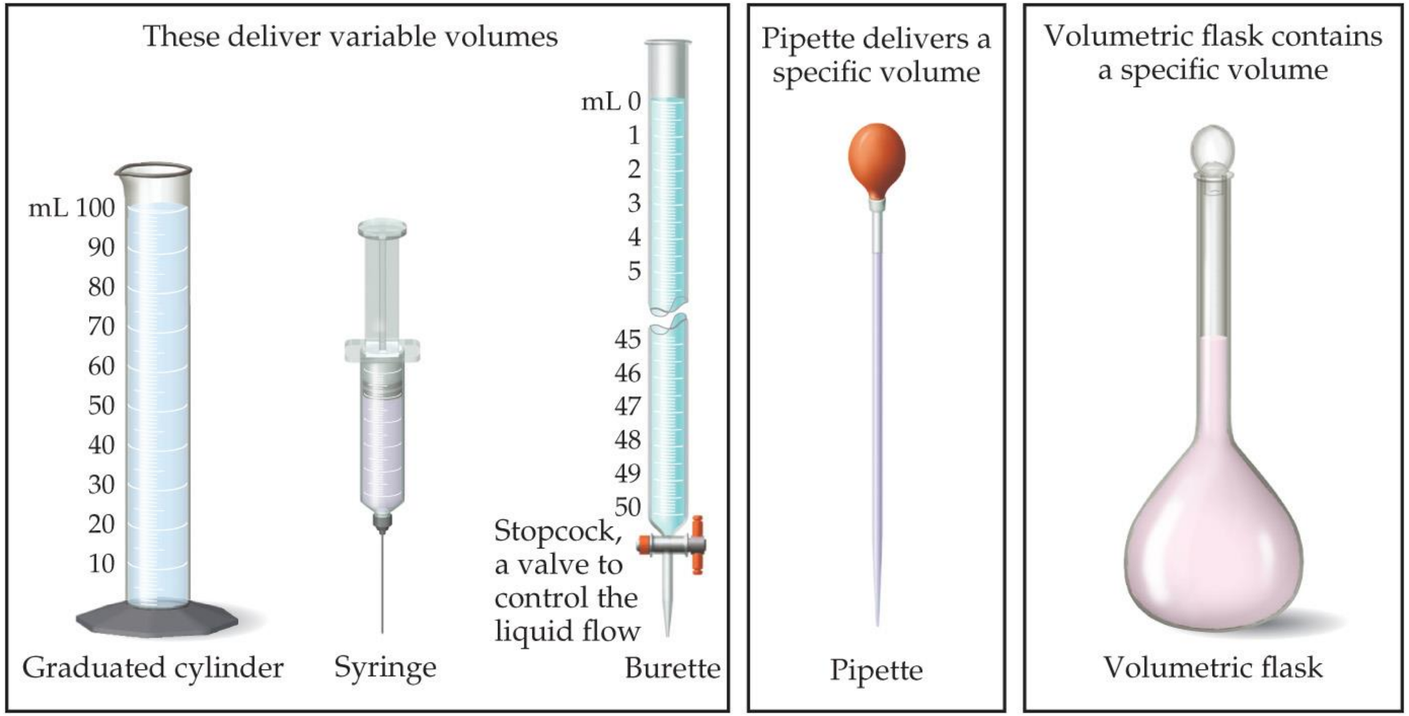


Density

- Density is a physical property of a substance.
- It has units that are derived from the units for mass and volume.
- The most common units are g/mL or g/cm^3.
- $d = m/V$

7.Uncertainities in Measurement

- All measured numbers have some degree of inaccuracy.
- Different measuring devices have different uses and different degrees of accuracy.



Uncertainty in Measurement: Exact Numbers

Numbers encountered in science:

- **Exact numbers:** Counted or given by definition.
- **Inexact numbers:** Measured (with uncertanity)

Examples:

- There are 12 eggs in 1 dozen.
- 1000 g in a kg
- 2.54 cm in an inch
- 1m = 100 cm
- 1kg = 2.2046 lb
- Counting objects: exact number of people in classroom

Uncertainty in Measurement: Inexact Numbers

Numbers obtained by measurement are always inexact. **Inexact** (or **measured**) numbers depend on how they were determined.

Equipment errors & Human errors !

10 students measure a dime using 10 balances

Calibration of the instruments could be different

Different students can read different masses

Uncertainties always exist in measured quantities

Give it some thought

Which of the following is an **inexact quantity?**

a) The number of people in chemistry class
b) The mass of a penny
c) The number of grams in a kilogram

Accuracy and Precision

The terms precision and uncertainty are often used in discussing the uncertainties of measured values.

- **Accuracy** refers to the proximity of a measurement to the true value of a quantity.
- **Precision** refers to the proximity of several measurements to each other.

Good accuracy
Good precision

Poor accuracy
Good precision

Poor accuracy
Poor precision

8.Significant Figures

- The term **significant figures** refers to digits that were measured.
- Mass of a dime on a balance: 2.2405 ± 0.0001 g. there is always uncertainty in the last digit reported
- The ± notation expresses the uncertainty in a measurement
- All digits of measured quantity including the uncertain one are called significant figures

8.Significant Figures

- 2.4 g → 2 significant figures
  2.2405 g → 5 significant figures
- What difference exist between measured values of **4.0 g** and **4.00** g

4.0: 2 significant figures • more uncertainty • in the first decimal place • the mass is closer to 4.0 than 3.9 or 4.1 • 4.0 ± 0.1 g

4.00: 3 significant figures • less uncertainty • in the second decimal place • the mass is closer to 4.00 than 3.99 or 4.01 g • 4.00 ± 0.01 g

8.Significant Figures

To determine the number of significant figures in a reported measurement:

- Read the number from left to right
- Counting the digits starting with the first digit that is not zero
- In any measurement that is properly reported all non zero digits are significant

00<u>12345</u>

8.Significant Figures

Zeros may or may not be significant:

Zeros are used : as a a part of the measured value

or

to locate the decimal point

1.

Zeroes between two significant figures are themselves **significant**

Ex:

1005 kg: 4 significant figures

7.03 cm: 3 significant figures

8.Significant Figures

2.

Zeroes at the beginning of a number are **never significant**, they indicate the position of the decimal point:

0.0<u>2</u> g, 0.00<u>26</u> cm

3.

Zeros at the end of the number are significant if the number contains a decimal point:

0.0<u>200</u> g, <u>3.0</u> cm

8.Significant Figures

When a number ends with zeros but contains no decimal point:

- The zeros are assumed not significant
- **Exponential notation** can be used to indicate **whether or not** end zeros are significant

Ex: 10,300 can be 3, 4, 5 significant figures

$\underline{1.03} \times 10^{4}$ g
$\underline{1.030} \times 10^{4}$ g
$\underline{1.0300} \times 10^{4}$ g

9.Significant Figures in Calculations

1. **Addition and substraction:** answers are rounded to the **least decimal place**. The result has the same number of decimal places as the measurement with the fewest decimal places

20.4<u>2</u>	2 decimal place
1.32<u>2</u>	3 decimal place
83.<u>1</u>	**1 decimal place & limiting**
104.<u>842</u>	**round to 1 decimal place: 104.8**

2. When multiplication or division is performed, answers are rounded to the number of digits that corresponds to the ***least* number of significant figures** in any of the numbers used in the calculation. The result contains the same number of significant figures as the measurement with the fewest significant figures.

Area: (6.221 cm) (5.2 cm) = 32.3492 cm^2

round off to 32 cm^2

because 5.2 has 2 significant figures

10.Rounding of Numbers

In rounding of numbers look at the leftmost digit to be removed

- If the leftmost digit removed is less than 5, the preceding number does not change. Thus rounding off 7.248 to two significant figures gives 7.2

- If the leftmost digit removed is greater or equal, the preceding number is increased by 1. Rounding of 4.735 to the three significant figures gives 4.74 and rounding off 2.376 to two significant figures gives 2.4

11.Dimensional Analysis

- Keep track of units for measured numbers
- We use **dimensional analysis** to convert one quantity to another to obtain proper unit
- Most commonly, dimensional analysis utilizes **conversion factors** (e.g., 1 in. = 2.54 cm).
- We can set up a ratio of comparison for the equality either 1 in/2.54 cm *or* 2.54 cm/1 in.
- We use the ratio which allows us to change units (puts the units we have in the denominator to cancel).
- Equivalent units cancel each other

Dimensional Analysis: Example

- Example: 2.54 cm and 1 in. are the same length

$$\frac{2.54\ cm}{1\ in.} \qquad \frac{1\ in.}{2.54\ cm}$$

$$Number\ of\ centimeters = \underbrace{8.5\ \cancel{in.}}_{given\ unit} \times \frac{2.54\ cm}{1\ \cancel{in.}} = \underbrace{21.6\ cm}_{desired\ unit}$$

$$\boldsymbol{\cancel{given}\ unit\ x\ \frac{desired\ unit}{\cancel{given}\ unit} = desired\ unit}$$

Examine units. What is the given unit? What is the desired unit? What conversion factors are avaliable?

Converting units

Sample exercise 1.10.

If a women has a mass of 115 lb what is her mass in grams?
(1lb = 453.59 g)

$$Mass\ in\ grams = (115\ \cancel{lb}) \times \left(\frac{453.6\ g}{1\ \cancel{lb}}\right)$$

$$= 52{,}162\ g$$

$$= 5.22\ x\ 104\ \ g$$

Round off: Least signifincat figures

Practice Exercises

- Practice Exercise 1. Using SI prefixes

Which of the following weights would you expect to be suitable for weighing on an ordinary bathroom scale?

(a) 2.0×10^{7} mg

(b) 2500 μg

(c) 5×10^{-4} kg

(d) 4×10^{6} cg

(e) 5.5×10^{8} dg

Solution to Practice Exercise 1. Using SI prefixes

The weight of an average human being:

30 kg – 150 kg, so we must convert the units given to kg

(a) $2.0 \text{ x}10^7$ mg

$$2.0 \text{ x}10^7 \text{ mg} \xrightarrow{\text{x}10^{-3}} 2.0 \text{ x}10^4 \text{ g} \xrightarrow{\text{x}10^{-3}} 2.0 \text{ x}10^1 \text{ kg}$$

		kg	
$\text{x}10^3$	:1000	g	x 1000
$\text{x } 10^{-1}$	: 10	dg	x 10
$\text{x } 10^{-1}$	: 10	cg	x10
$\text{x } 10^{-1}$	: 10	mg	x10

Practice Exercise 2. Using SI prefixes

(a) How many picometers are there in 1 m?

m	
	x10
dm	
	x10
cm	
	x10
mm	
	x1000
µm	
	x1000
nm	
	x1000
pm	

$1m = 10^{12}$ pm

Practice Exercise 2. Using SI prefixes

(b) Express 6.0×10^3 m using a prefix to replace the power of ten. 6.0 km

(c) Use exponential notation to express 4.22 mg in grams

1 mg = 10^{-3} g   mg has a prefix of 10^{-3}

g

x10

dg

x10

cg

x10

mg

4.22×10^{-3} g

(d) Use decimal notation to express 4.22 mg in gram

g	
	x10
dg	
	x10
cg	
	x10
mg	

4.22 mg :1000 = 0.00422 g

Determining density and using density to determine volume or mass. Sample Exercise 1.4.

(a) Calculate the density of a mercury if 1.00 x 10^2 g occupies a volume of 7.36 cm^3

$$denisty = \frac{mass}{volume} = \frac{1.00x102\ g}{7.36\ cm^3} = 13.6\ g/cm^3$$

(b) calculate the volume of 65.0 g liquid methanol if its density is 0.791 g/ml

$$denisty = \frac{mass}{volume} \longrightarrow volume = \frac{mass}{density}$$

$$volume = \frac{65.0\ g}{0.791\ g/\cancel{ml}} \longrightarrow volume = 82.2\ ml$$

3 significant figures ! Multiplication and division: least significant figures !

Relating Significant Figures to the uncertainty of a measurement. Practice exercise 2.

A sample that has a mass of about 25 g is weighed on balance that has a precision of ± 0.001 g. How many significant figures should be reported for this measurement?

Significant figure would be 5 as in the measurements :

25.001 & 24.499

Assigning appropriate significant figures. Practice exercise 1.

Which of the following numbers in your personal life are exact numbers?

(a) Your phone number
(b) Your weight
(c) Your IQ
(d) Your driver's license number
(e) The distance you walked yesterday

Determining the number of significant figures in a measurement.
Sample exercise 1.7.

How many significant figures are in each of the following numbers (assume each number is a measured quantity)? (a) 4.003 (b) 6.023×10^{23}.

(a) 4.003 has 4 significant figures: all digits and zeros between numbers are considered significant
(b) 6.023 has 4 significant figures due to the reasons apply (a) and exponential numbers does not add as significant figure

Determining the number of significant figures in a measurement. Practice exercise 2.

How many significant figures are in each of the following measurements? (a) 3.549 g (b) 2.3×10^4 cm (c) 0.00134 m^3

(a) 3.549 g : 4 significant figures
(b) 2.3×10^4 cm : 2 significant figures
(c) 0.00134 m^3 : 3 significant figures. We do not count the zeros before a number, they show a decimal place

Uncertainty in Measurement. Exercise 1.37

What is the significant figures in each of the following measured quantities?

(a) 510 kg
(b) 0.0631 s
(c) 4.2604 cm
(d) 0.000479 L
(e) 7.05000 x 10^{-6} m^3
(f) 800 m

Remember the rules!

(a) 2 significant figures, do not count the zeros at the end unless there is decimal point
(b) 3 significant figures dot count the zeros before a number
(c) 5 significant figures: count zeros between numbers
(d) 3 significant figures: do not count the numbers before digits
(e) 6 significant figures: exponential number is used to show significant figures
(f) 1 significant figure, do not count the zeros at the end unless there is decimal point

Uncertainty in Measurement. Exercise 1.39

Round each of the following numbers to **four significant figures** and express the result in standard **exponential notation**

(a)102.53070

(b) 656.980

(c) 0.008543210

(d) 0.000257870

(e) -0.0357202

Check if the leftmost digit is >= 5 or not

(a) 102.53070 : 102.5 : 1.205×10^{3}

(b) 656.980 : 657.0 : 6.570×10^{2}

(c) 0.008543210 : 0.008543 : 8.543×10^{-3}

(d) 0.000257870 : 0.0002579 : 2.579×10^{-4}

(e) -0.0357202 : -0.03572 : 3.572×10^{-2}

Uncertainty in Measurement. Exercise 1.41

Carry out the following operations and express the answers with the appropriate number of significant figures.

(a) 25.3693 + 1.78

(b) 834.5 – 212.75648

(c) 128.56 x 3186.14862

(d) 0.0482 / 0.941

Rules:

1. Addition and subtraction: least decimal place
2. Multiplication and division: least significant figures

(a) 25.3693 + 1.78 = 27.1493 is rounded to 27.15

Rule 1. 4 decimal place and 2 decimal places so the answer must have 3 decimal places and should be rounded properly

(b) 834.5 - 212.75648 = 621.74352 = 621.7

Rule 1. 1 and 5 decimal points, so the answer must have 1 decimal point

(c) 128.56 x 3186.14862 = 409611,267 is rounded of to 409610 with 5 significant figures

Rule 2. 5 and 9 significant figures, so the answer must contain 5 significant figures

(d) 0.0482 / 0.941= 0.0512221 is rounded to 0.0512

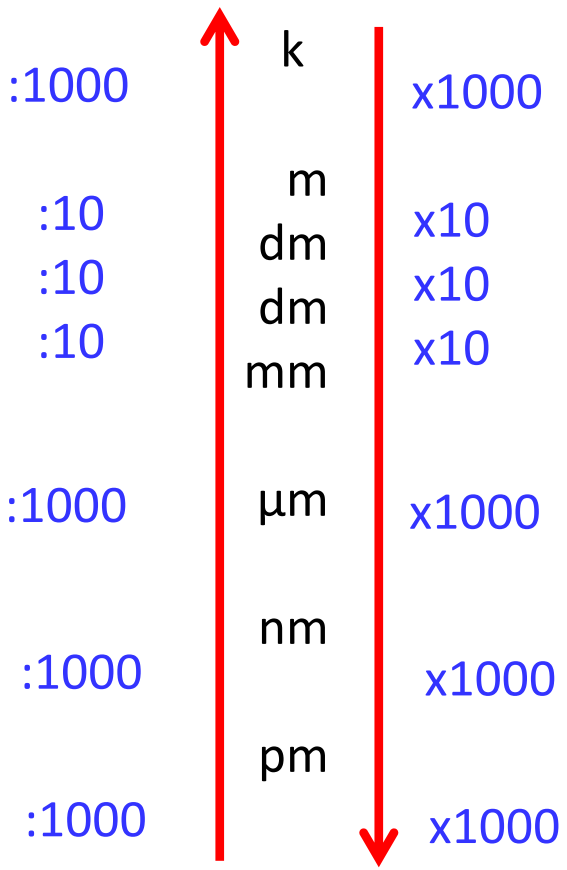
Rule 2. 3 significant figures, so the answer must contain 3 significant figures

Dimensional Analysis Exercise 1.45

Using knowledge of metric units, English units and information on the back inside cover, write down the conversion factors needed to convert

(a) mm to nm
(b) mg to kg
(c) km to ft
(d) in^3 to cm^3

(a) $1m = 10^9$ nm
(b) $1\ mg = 0.001\ g = 0.000001\ kg = 10^{-6}\ kg$



(c) 1 km = 0.62137 mi

1 mi = 5280 ft

given in back inside cover of the book

To convert ft to mi

$$\boxed{desired\ unit = given\ unit\ x\ \frac{desired\ unit}{given\ unit}}$$

$$\boxed{ft = 0.62137\ mi\ x\ \frac{5280\ ft}{mi}}$$

$$\boxed{= 3280.83\ \text{ft}}$$

3280.83 rounded to 3280 with 3 significant figures

(d) in^3 to cm^3

1 in. = 2.54 cm

1 in^3 = 2.54^3 cm

1 in^3 = 41.6 cm^3 (3 significant figures)

Dimensional Analysis Exercise 1.47

(a) A bumblebee flies with a ground speed of 15.2 m/s. Calculate its speed in miles per hour

1 min = 60 s
1 hr = 60 min

Conversion factors between hr and s:

$$\frac{1\ min}{60\ s} \times \frac{1\ hr}{60\ min} = \frac{1\ hr}{3600\ s} \quad \text{or} \quad \frac{0.0028\ hr}{1\ s} \quad \text{or} \quad \frac{1\ s}{0.00028\ hr}$$

1 km = 0.62137 mi and 1 km = 1000 m

Conversion factors :

$$\frac{1\ km}{0.62137\ mi} = \frac{1000\ m}{0.62137\ mi} = \frac{1\ m}{0.0062137\ mi}$$

So we plug the numbers in m/s

$$\frac{m}{s} = \frac{0.0062137\, mi}{0.00028\, hr}$$

1 m/s = 2.2 mi/hr

15.2 m/s = 2.2 x 15.2 m/hr

= 33.4 mi/hr

Classification and Properties of Matter

1.13. Classify each of the following as a pure substance or a mixture

(a) Calcium
(b) Lake water
(c) Chocolate pudding
(d) Cruncy peanut butter

1.19. In the process of attempting to characterize a substance, a chemist makes the following observations: the substance is silvery white, lustrous metal. It melts at 649 ^{0}C and boils at 1105 ^{0}C. Its density at 20 ^{0}C is 1.738 g/cm^3. The substance burns in air, producing an intense white light. It reacts with chlorine to give a brittle white solid. The substance pounded into thin sheets or drawn into wires. It is good conductor of electricity. Which of these chemical characteristics are physical properties, and which are chemical properties?

1.21. Label each of the following as either a physical process or a chemical process:

(a) boiling a pot of water

(b) digesting a meal

(c) converting grain into flour

(d) exploding of nitrous glycerine

(e) rusting of a nail

1.61. A sample of ascorbic acid, i.e. Vitamin C, is synthesized in the laboratory. It contains 1.50 g of carbon and 2.00 g of oxygen. Another sample of ascorbic acid isolated from citrus fruits contains 6.35 g of carbon. How many grams of oxygen does it contain? Which law are you assuming in answering this question.

Total mass of ascorbic acid synthesized in lab = 1.50 g + 2.00 g = 3.50 g

$$\frac{mass\ of\ carbon}{mass\ of\ ascorbic\ acid} = \frac{1.50\ g}{3.50\ g}$$

According to the ***law of constant composition***, atoms combined in fixed ratio

$$\frac{mass\ of\ carbon}{6.35\ g} \cancel{=} \frac{1.50\ g}{3.50\ g}$$

mass of carbon = 2.72 g

Sample Exercise 1.2. Using SI prefixes

What is the name of the unit equals to

(a) 10^{-9} gram
(b) 10^{-6} seconds
(c) 10^{-3} meter


thank you
danke
謝謝
ngiyabonga
teşekkür ederim
спасибо
gracias
dank je
tapadh leat
mochchakkeram
dziękuję
go raibh maith agat
obrigado
sukriya
kop khun krap
arigato
dakujem
grazie
мерси
merci
감사합니다
bedankt
hvala